

## 10.1: Introduction to Acids and Bases

### Learning Objectives

Make sure you thoroughly understand the following essential ideas which have been presented above.

- Suggest simple tests you could carry out to determine if an unknown substance is an acid or a base.
- State the chemical definitions of an acid and a base in terms of their behavior in water.
- Write the formula of the salt formed when a given acid and base are combined.

The concepts of an acid, a base, and a salt are very old ones that have undergone several major refinements as chemical science has evolved. Our treatment of the subject at this stage will be mainly conceptual and qualitative, emphasizing the definitions and fundamental ideas associated with acids and bases. We will not cover calculations involving acid-base equilibria in these lessons.

### 1 Acids

The term *acid* was first used in the seventeenth century; it comes from the Latin root *ac-*, meaning “sharp”, as in *acetum*, vinegar. Some early writers suggested that acidic molecules might have sharp corners or spine-like projections that irritate the tongue or skin. Acids have long been recognized as a distinctive class of compounds whose aqueous solutions exhibit the following properties:

- A characteristic sour taste (think of lemon juice!);
- ability to change the color of litmus from blue to red;
- react with certain *metals* to produce gaseous  $H_2$ ;
- react with *bases* to form a salt and water.

**Litmus** is a natural dye found in certain lichens. The name is of Scandinavian origin, e.g. *lit* (color) + *mosi* (moss) in Icelandic. “Litmus test” has acquired a meaning that transcends both Chemistry and science to denote any kind of test giving a yes/no answer.

#### How oxygen got mis-named

The first chemistry-based definition of an acid turned out to be wrong: in 1787, Antoine Lavoisier, as part of his masterful classification of substances, identified the known acids as a separate group of the “complex substances” (compounds). Their special nature, he postulated, derived from the presence of some common element that embodies the “acidity” principle, which he named *oxygen*, derived from the Greek for “acid former”.

Lavoisier had recently assigned this name to the new gaseous element that Joseph Priestly had discovered a few years earlier as the essential substance that supports combustion. Many combustion products (oxides) do give acidic solutions, and oxygen is in fact present in most acids, so Lavoisier’s mistake is understandable. In 1811 Humphrey Davy showed that muriatic (hydrochloric) acid (which Lavoisier had regarded as an element) does not contain oxygen, but this merely convinced some that chlorine was not an element but an oxygen-containing compound. Although a dozen oxygen-free acids had been discovered by 1830, it was not until about 1840 that the hydrogen theory of acids became generally accepted. By this time, the misnomer *oxygen* was too well established a name to be changed. The root *oxy* comes from the Greek word  $\acute{o}\xi\upsilon\varsigma$ , which means “sour”.

### 2 Acids and the hydrogen ion

The key to understanding acids (as well as bases and salts) had to await Michael Faraday’s mid-nineteenth century discovery that solutions of salts (known as electrolytes) conduct electricity. This implies the existence of charged particles that can migrate under the influence of an electric field. Faraday named these particles *ions* (“wanderers”). Later studies on electrolytic solutions suggested that the properties we associate with acids are due to the presence of an excess of *hydrogen ions* in the solution. By 1890 the Swedish chemist Svante Arrhenius (1859-1927) was able to formulate the first useful theory of acids:

"an acidic substance is one whose molecular unit contains at least one hydrogen atom that can dissociate, or *ionize*, when dissolved in water, producing a hydrated hydrogen ion and an anion."

hydrochloric acid	$\text{HCl} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
sulfuric acid	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq})$
hydrogen sulfite ion	$\text{HSO}_3^-(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$
acetic acid	$\text{H}_3\text{CCOOH} \rightarrow \text{H}^+(\text{aq}) + \text{H}_3\text{CCOO}^-(\text{aq})$

Strictly speaking, an "Arrhenius acid" must contain hydrogen. However, there are substances that do not themselves contain hydrogen, but still yield hydrogen ions when dissolved in water; the hydrogen ions come from the water itself, by reaction with the substance.

An acid is a substance that yields an excess of hydrogen ions when dissolved in water.

There are three important points to understand about hydrogen in acids:

- Although all Arrhenius acids contain hydrogen, not all hydrogen atoms in a substance are capable of dissociating; thus the  $\text{-CH}_3$  hydrogens of acetic acid are "non-acidic". An important part of knowing chemistry is being able to predict which hydrogen atoms in a substance will be able to dissociate into hydrogen ions.
- Those hydrogens that do dissociate can do so to different degrees. The *strong acids* such as HCl and  $\text{HNO}_3$  are effectively 100% dissociated in solution. Most organic acids, such as acetic acid, are *weak*; only a small fraction of the acid is dissociated in most solutions. HF and HCN are examples of weak inorganic acids.
- Acids that possess more than one dissociable hydrogen atom are known as *polyprotic acids*;  $\text{H}_2\text{SO}_4$  and  $\text{H}_3\text{PO}_4$  are well-known examples. Intermediate forms such as  $\text{HPO}_4^{2-}$ , being capable of both accepting and losing protons, are called *ampholytes*.

$\text{H}_2\text{SO}_4$ sulfuric acid	→	$\text{HSO}_4^-$ hydrogen sulfate ("bisulfate") ion	→	$\text{SO}_4^{2-}$ sulfate ion	
$\text{H}_2\text{S}$ hydrosulfuric acid	→	$\text{HS}^-$ hydrosulfide ion	→	$\text{S}^{2-}$ sulfide ion	
$\text{H}_3\text{PO}_4$ phosphoric acid	→	$\text{H}_2\text{PO}_4^-$ dihydrogen phosphate ion	→	$\text{HPO}_4^{2-}$ hydrogen phosphate ion	→ $\text{PO}_4^{3-}$ phosphate ion
$\text{HOOC-COOH}$ oxalic acid	→	$\text{HOOC-COO}^-$ hydrogen oxalate ion	→	$^- \text{OOC-COO}^-$ oxalate ion	

You will find out in a later section of this lesson that hydrogen ions cannot exist as such in water, but don't panic! It turns out that chemists still find it convenient to pretend as if they are present, and to write reactions that include them.

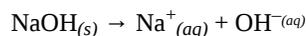
### 3 Bases

The name *base* has long been associated with a class of compounds whose aqueous solutions are characterized by:

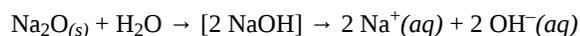
- a bitter taste;
- a "soapy" feeling when applied to the skin
- ability to restore the original blue color of litmus that has been turned red by acids
- ability to react with acids to form salts.
- react with certain *metals* to produce gaseous  $\text{H}_2$

The word “alkali” is often applied to strong inorganic bases. It is of Arabic origin, from *al-kali* (“the ashes”) which refers to the calcined wood ashes that were boiled with water to obtain *potash* which contains the strong base KOH, used in soap making. The element name *potassium* and its symbol K (from the Latin *kalium*) derive from these sources.

Just as an acid is a substance that liberates hydrogen ions into solution, a *base* yields *hydroxide ions* when dissolved in water:



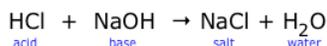
Sodium hydroxide is an *Arrhenius base* because it contains hydroxide ions. However, other substances which do not contain hydroxide ions can nevertheless produce them by reaction with water, and are therefore also classified as bases. Two classes of such substances are the metal oxides and the hydrogen compounds of certain nonmetals:



A base is a substance that yields an excess of hydroxide ions when dissolved in water.

## 4 Neutralization

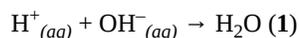
Acids and bases react with one another to yield two products: water, and an ionic compound known as a *salt*. This kind of reaction is called a *neutralization reaction*.



This "molecular" equation is convenient to write, but we need to re-cast it as a net ionic equation to reveal what is really going on here when the reaction takes place in water, as is almost always the case.



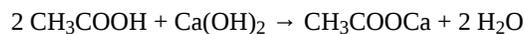
If we cancel out the ions that appear on both sides (and therefore don't really participate in the reaction), we are left with the net equation



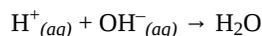
which is the fundamental process that occurs in all neutralization reactions.

Confirmation that this equation describes all neutralization reactions that take place in water is provided by experiments indicating that no matter what acid and base are combined, all liberate the same amount of heat (57.7 kJ) per mole of H<sup>+</sup> neutralized.

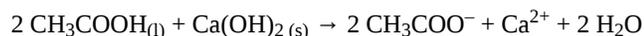
In the case of a weak acid, or a base that is not very soluble in water, more than one step might be required. For example, a similar reaction can occur between acetic acid and calcium hydroxide to produce *calcium acetate*:



If this takes place in aqueous solution, the reaction is really between the very small quantities of H<sup>+</sup> and OH<sup>-</sup> resulting from the dissociation of the acid and the dissolution of the base, so the reaction is identical with Equation 1:



If, on the other hand, we add *solid* calcium hydroxide to pure liquid acetic acid, the net reaction would include both reactants in their "molecular" forms:



The “salt” that is produced in a neutralization reaction consists simply of the anion and cation that were already present. The salt can be recovered as a solid by evaporating the water.

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